Chapter-2

Lesson-1

I - Atomic Theory

A) Postulates of Early Atomic Theory:

1) Elements are composed of tiny particles called atoms.
2) Atoms of the same element are exactly alike and have the same set of properties, but the atoms of one element are different from the atoms of different elements.
3) Atoms of an element cannot be created, destroyed, or changed into atoms of another element.
4) Compounds are formed when atoms of two or more elements are chemically combined in a fixed small whole number ratio.

B) Evidence to Support Early Atomic Theory:

- During a chemical reaction there is no detectable change between the total mass of the reactants and the total mass of the products.
- (Law of Conservation of Mass)
  - **A synthesis reaction between element-A and element-B**

    1.188 g of A and 0.711 g of B → 1.899 g of compound AB

    1.000 g of A and 0.598 g of B → 1.598 g of compound AB

- A Compound is always composed of the same elements in the same ratio (percentage) by mass.
  (Law of Constant Composition)
  - **A compound, AB, consists of element-A and element-B**

    1.188 g of A and 0.711 g of B → 62.6% A and 37.4 % B

    1.000 g of A and 0.598 g of B → 62.6% A and 37.4 % B

- When two elements form more than one compound and the mass of one of the elements in each compound is held constant at 1.0 g, then the masses of the second element will be found to occur in an exact ratio of small whole numbers to each other.
  (Law of Multiple Proportions)

Element-A and element-B will combine to form two different compounds:

Compound-1	Compound-2
1.188 g of A	0.396 g of A
0.711 g of B	0.474 g of B

making the mass of element-A 1.000 g in both compounds, the mass ratios become

Compound-1	Compound-2
1.188 g of A ÷ 1.188 = 1.000 g of A	0.396 g of A ÷ 0.396 = 1.000 g of A
0.711 g ÷ 1.188 of B = 0.598 g of B	0.474 g of B ÷ 0.396 = 1.20 g of B

The mass ratio of element-B in each compound is

Compound-1		Compound-2
0.598 g of B ÷ 0.598 g		1.20 g of B ÷ 0.598 g
$\boxed{\text{1 of B}}$	**1 to 2**	$\boxed{\text{2 of B}}$

Q1: Does the following data illustrate the Law of Multiple Proportions?

Compound-1	Compound-2
16.684 g of Sulfur	4.005 g of Sulfur
16.649 g of Oxygen	5.995 g of Oxygen

A1: Make the mass of sulfur 1.000 g in both compounds.

Compound-1	Compound-2
1.000 g of Sulfur	1.000 g of Sulfur
0.998 g of Oxygen	1.497 g of Oxygen

Calculate the mass ratio of oxygen in each compound.

0.998 g of O ÷ 0.998		1.497 g of O ÷ 0.998
1 g of O		1.512 g of O → [0.512 ≈ ½]
1 g × 2 = $\boxed{2}$	**2 to 3**	1.512 g × 2 ≈ $\boxed{3}$

YES the data illustrates Law of Multiple Proportions.

II – Atomic Structure

A) Subatomic Particles:

Particle	Relative mass	Relative Charge	Location
proton	1 μ	+1	nucleus
neutron	1 μ	0	nucleus
electron	≈0 μ (*0.00055*μ)	-1	outside nucleus

- The Cathode Ray Tube Experiments (J.J. Thomson)
  - Found that elements possessed small negatively charged particles (electrons).
  - Found that the mass-to-charge ratio of these electrons was 5.68×10^{-9} g/coulomb.
- The Oil Drop Experiment (R. Millikan)
  - Found that the charge of an electron was 1.60×10^{-19} coulomb.
  - Using the mass-to-charge ratio of an electron, found that its mass was 9.10×10^{-28} g.
- Radioactivity Experiments (E. Rutherford)
  - Found three types of radiation (α, β, and γ).

Particle	Alpha (α).	Beta (β).	Gamma (γ).
Mass	4 μ	≈0 μ (*0.00055*)	0 μ
Charge	+2	-1	0

- The Gold Foil Experiment (E. Rutherford)
  - Found that an atom's volume is mostly empty space.
  - Found that most of an atom's mass is contained in an extremely small positive charged central volume called the nucleus.
- The Visible Light Emission Spectra Experiments (J. Balmer et al.)
  - Found that elements had a unique set of spectral lines.
  - Each line was determined to represent an electron energy "jump" between energy levels.
- The X-ray Emission Spectra Experiment (H. Moseley)
  - Found that a plot of the square root of the elements' X-ray emission frequency versus the order number of the elements was a perfect straight line.
  - The order number became the atomic number.

B) Atomic Number:

- The number of protons in a nucleus is called the ***atomic number***.
- In a neutral atom, the number of electrons is equal to the number of protons.

C) Mass Number & Isotopes:

- The sum of an atom's number of protons and neutrons is called the ***mass number***. The mass number is always an integer!
- When two atoms of the same element have a different mass number we say that these atoms are ***isotopes*** of the same element.
- ***Isotopes*** of the same element must have the same number of protons but then have different numbers of neutrons. Thus isotopes are known by their mass numbers.
- An atom of a specific isotope is called a ***nuclide***.
  - ^{12}C (6p ; 6n) & ^{14}C (6p ; 8n)
  - Cl-35 (17p ; 18n) & Cl-37 (17p ; 20n)

III – Atomic Weights

A) The Atomic Mass Scale:

- Since 1961, an atom of C-12 has been assigned the atomic mass of exactly 12 amu (μ). All other atoms have been compared to this standard.
- Thus, 1 amu = 1.66054×10^{-24} g and 1 g = 6.02214×10^{23}amu
- The masses of other isotopes are determined by finding the relative mass of the isotope in question to C-12.
- The relative mass of an isotope is equal to the ratio of the voltage needed to bring the isotope to a point in a mass spectrometer to the voltage needed to bring C-12 to that same point in the mass spectrometer. [***Read" A Closer Look" on page 48 in textbook***]

Q2: What is the mass of O-16 if its relative mass to C-12 is 1.33291?

A2: $$\frac{\text{mass } ^{16}\text{O atom}}{\text{mass } ^{12}\text{C atom}} = 1.33291 \rightarrow 1.33291\times12 = \boxed{15.9949 \text{ amu}}$$

Q3: What is the mass of ^{13}C if its relative mass to ^{12}C is 1.0836129?

A3: $\dfrac{\text{mass } ^{13}\text{C atom}}{\text{mass } ^{12}\text{C atom}} = 1.0836129 \rightarrow 1.0836129 \times 12 = \boxed{13.003355\ \mu}$

B) Average Atomic Masses and Isotopic Abundances:

- Atomic masses are weighted averages of the various isotopes of the element.
- To calculate a weighted average one must know:
  - the isotopic mass of each isotope (*or use the mass numbers as approximate masses in the absence of isotopic masses*).
  - the isotopic abundance of each isotope. (*See the summer assignment*)

Q4: What is the weighted average (atomic weight) of C-12 (98.93%), C-13 (13.003355 amu and 1.07%)?

A4: $(12)(0.9893) + (13.003355\ \mu)(0.0107)$

$(11.87) + (0.139) = \boxed{12.01\text{amu}}$

Q5: What is the weighted average (atomic weight) of Ne-20 (90.92%), Ne-21 (0.26%), and Ne-22 (8.82%)?

A5: $(\approx 20)(0.9092) + (\approx 21)(0.0026) + (\approx 22)(0.0882)$

$(18.18) + (0.055) + (1.94) = \boxed{\approx 20.18\mu}$

Q6: What are the percent abundances of ^{10}B (10.01μ) and ^{11}B (11.01μ) if boron's atomic mass is 10.81μ

A6: $(X)(10.01) + (1-X)(11.01) = 10.81$

$11.01 - 1X = 10.81$

$X = 0.2 \rightarrow \boxed{20\%\ ^{10}B}$

$1-X = 0.8 \rightarrow \boxed{80\%\ ^{11}B}$

C) Early Attempts to Determine Atomic Weights:

- Dalton prepared the first table of atomic weights (1808). Many of his values were found to be incorrect because of flawed assumptions about the formulas of certain compounds. ***Dalton knew that water contained 8 g of oxygen for every 1 g of hydrogen. He concluded that the formula for water was OH. Arbitrarily assigning hydrogen a mass of 1, oxygen's relative mass became 8***.

- An element's specific heat is the amount of heat needed to raise the temperature of 1g of the element by 1^{o}C and this amount of heat is dependent on the type of atom. Dulong and Petit suggested that the amount of heat needed to raise the temperature of a mole of atoms of a solid element by 1^{o}C should be independent of the type of atom. They concluded that (1819):

  (molar mass) × (specific heat) ≈ 25 J/mol^{o}C

  (MM of Ag) × (0.236) ≈ 25 J/mol^{o}C

  (MM of Ag) ≈ 106 g/mol **(accepted value is ≈ 108 g/mol)**

- Based on work by Gay-Lussac, Avogadro and Richter on the proportions in which substances combine, as well as a fanatical dedication to observation and measurement, Berzelius was able to publish (1826) the first highly accurate table of atomic masses of the elements.

Element	Dalton's Atomic Masses	Berzelius' Atomic Masses
Copper	56	63.00
Hydrogen	1	1.00
Lead	95	207.12
Oxygen	8	16.00
Silver	100	108.12
Sulfur	13	32.18

- Mendeleev's Periodic Table (1869) was an additional tool for estimating atomic masses. The atomic mass of an element should be approximately equal to the average of the atomic masses of the elements to the right and left of its position on the table. MM of Sc ≈ 45… [(MM of Ca = 40 + MM of Ti = 48) ÷ 2 = 44]

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IV – The Periodic Table

A) Structure:

- Elements are arranged in increasing order of their atomic masses.
- Elements with similar properties are placed in the same column in the table. These columns are called **groups**.
  - There are 18 groups that are labeled group 1 through 18.
  - The groups can also be labeled using an A and B designation.
    - The A-groups, called the "main group elements", are groups 1 and 2 and 13 to 18. Group 2A is group 2 and group 4A is group 14.
    - The B-groups are groups 3 through 12. The B group designations are a bit confusing.
- The horizontal rows of the table are called **periods**. There are 7 periods on the table and are labeled period 1 through 7.

B) Categorization:

- Except for hydrogen, the elements on the left side and middle of the table are **metals**. These elements share certain basic properties (high luster, high conductivity, high malleability, etc.).
- The elements at the upper right are **nonmetals**.
- The elements that border the metals and nonmetals are the **metalloids**. [usually: B, Si, Ge, As, Sb, Te, and At]
- Hydrogen is often labeled as a nonmetal.
- Specific groups on the table have names:
  - Group 1: alkali metals
  - Group 2: alkaline earth metals
  - Group 15: pnictides [*the suffocaters*]
  - Group 16: chalcogens [*the ore formers*]
  - Group 17: halogens [*the salts formers*]
  - Group 18: noble gases [*the inert gases; the rare gases*]

- Collective groups on the table have names:
  - Groups 1,2,13 – 18: main group elements
  - Groups 3 – 12: transition elements
  - Non-groups: inner transition elements [*rare earth elements*]

V – Molecules and Molecular Compounds

A) Molecules and Chemical Formulas:

- A notation that uses chemical symbols and numerical subscripts to convey the relative proportions of atoms of different elements in a substance is called a ***chemical formula***.
- A discrete grouping of tightly bonded atoms that possess the properties of a substance is called a ***molecule***.
- Most molecular substances contain only nonmetals.
- The bonding between atoms in most molecular substances is covalent.
- Individual noble gas atoms are considered to be molecules.
- Substances composed of molecules are called ***molecular substances***. *Examples*: He, N_2, O_3, H_2O, $C_6H_{12}O_6$
- Compounds composed of molecules are called ***molecular compounds***. *Examples*: CO_2, CH_3COOH,

B) Molecular and Empirical Formulas:

- The chemical formula for any molecule is called the ***molecular formula***. Molecular formulas indicate the actual number and types of atoms present in a molecule.
  - CH_2O, $C_6H_{12}O_6$, H_2O_2, HCl, CH_3NH_2, N_2
- The simplest chemical formula for any compound is called the ***empirical formula***.
  - K_2CrO_4, CH_2O, HgCl, HO, NaOH, $AlCl_3$

Q1: What are the empirical formulas for the following compounds?
a) glucose - $C_6H_{12}O_6$ b) acetic acid - $HC_2H_3O_2$
c) methanal - CH_2O d) methyl methanoate - $HCOOCH_3$

A1: a) CH_2O b) CH_2O c) CH_2O d) CH_2O

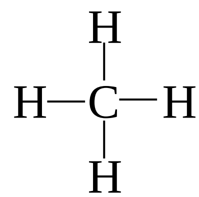
C) Picturing Molecules:

- A molecular formula does not show how the atoms are arranged and bonded in the molecule.
- A **structural formula** does show how the atoms are arranged and bonded in the molecule.

Example:

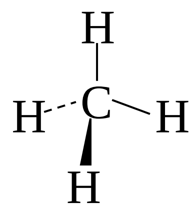
CH_4

molecular formula

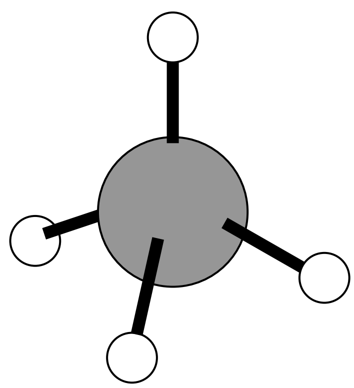


structural formula

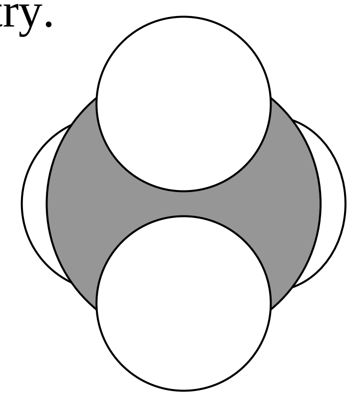
- A **perspective drawing** shows the geometry of the molecule.



- A **ball and stick model** shows the atoms as spheres and bonds as sticks as well as the geometry.



- A **space-filling model** shows the relative sizes of the atoms as well as the geometry.

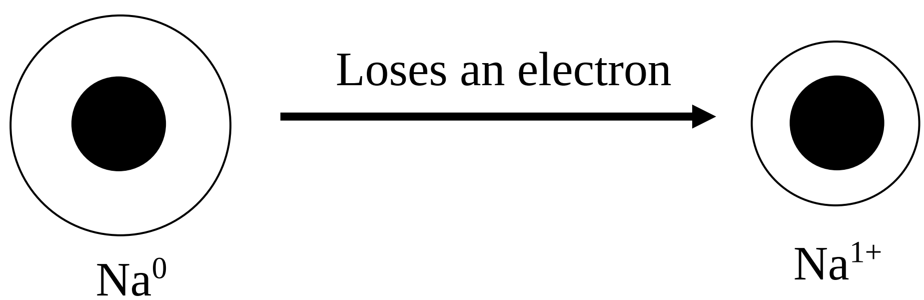


The space-filling model of a molecule is the most realistic even though the geometry is often difficult to see.

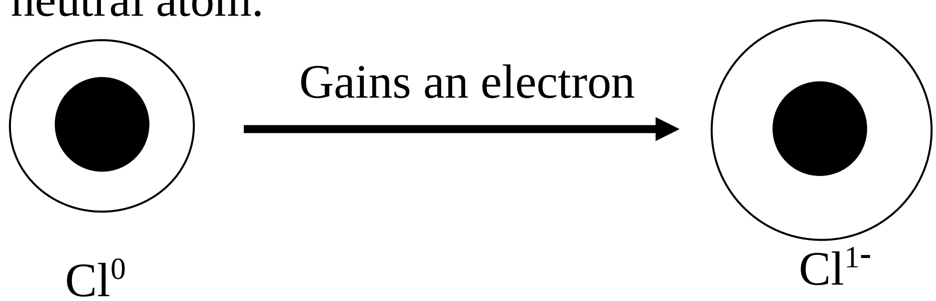
VI – Ions and Ionic Compounds

A) Ions:

- An atom is a neutral particle in which the number of protons in the nucleus is equal to the number of electrons outside the nucleus.
- If an atom gains or loses one or more electrons, an ***ion*** is formed. An ion is a charged particle.
- Ions have very different chemical properties from the chemical properties of their neutral atoms.
- A ***polyatomic ion*** is a group of bonded atoms that has a net charge.
- A positive ion is called a ***cation***. Cations are formed when an atom loses one or more electron. A cation is always smaller than its neutral atom.



- A negative ion is called a ***anion***. Anions are formed when an atom gains one or more electron. An anion is always larger than its neutral atom.



- Atoms of main group elements will gain or lose electron so as to end up with the same number of electrons as the noble gas nearest to them on the periodic table.

Q2: Predict the ionic charges of the following neutral atoms.
a) Li^0 b) Al^0 c) S^0 d) Br^0

A2: a) Li^{1+} b) Al^{3+} c) S^{2-} d) Br^{1-}

B) Ionic Compounds:

- Compounds composed of oppositely charged ions are called ***ionic compounds***. *Examples*: $NaCl$, $MgSO_4$, NH_4NO_3
- Simple ionic compounds consist of a metal and a nonmetal.
- The chemical formulas for most ionic compounds are written as empirical formulas.

Q3: Which of the following compounds would you expect to be ionic? HCl, NH_4Cl, NH_3, KH, $FeSO_4$, H_2SO_4

A3: NH_4Cl, KH, & $FeSO_4$

Q4: Write the chemical formula for the compound composed of Na^{1+} ions and CO_3^{2-} ions.

A4: Na_2CO_3

Q5: Write the chemical formula for the compound composed of Mg^{2+} ions and N^{3-} ions.

A5: Mg_3N_2

Q6: Write the chemical formula for the compound composed of Hg^{2+} ions and NO_3^{1-} ions.

A6: $Hg(NO_3)_2$

Q7: Why is Hg_2Cl_2 the correct formula for mercury (I) chloride and not $HgCl$?

A7: Because mercury (I) exists as a polyatomic ion. Hg_2^{2+}

Q8: Write the chemical formula for the compound composed of NH_4^{1+} ions and PO_3^{3-} ions.

A8: $(NH_4)_3PO_3$

Predict whether the compounds below are <u>ionic</u> or <u>molecular</u>.

Q9: NH_4NO_3

A9: ionic

Q10: HI

A10: molecular

Q11: KH

A11: ionic

Q12: $CaSO_4$

A12: ionic

Q13: SO_3

A13: molecular

Q14: SF_6

A14: molecular

Q15: FeO

A15: ionic

Q16: XeF_2

A16: molecular

Q17: $NaHCO_3$

A17: ionic

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VII – Naming Inorganic Compounds

A) Naming Inorganic Ionic Compounds:

1) Name the Cation:

a) Cations formed from metal atoms have the same name as the metal.

Na^{1+}	**sodium ion**	**Ag^{1+}**	**silver ion**

b) If the metal can form different cations, the positive charge is indicated by a Roman numeral in parentheses following the name of the metal.

Cu^{1+}	**copper (I)**	**Cu^{2+}**	**copper (II)**
Hg_2^{2+}	**mercury (I)****	**Hg^{2+}**	**mercury (II)**

An older method is still used for certain metals that can form different cations. The endings ***ic*** and ***ous*** are added to latin root names and are used to indicate the higher and lower charged ion, respectively.

Cu^{1+}	**cuprous**	**Cu^{2+}**	**cupric**
Fe^{2+}	**ferrous**	**Fe^{3+}**	**ferric**
Sn^{2+}	**stannous**	**Sn^{4+}**	**stannic**
Pb^{2+}	**plumbous**	**Pb^{4+}**	**plumbic**
Co^{2+}	**cobaltous**	**Co^{3+}**	**cobaltic**
Ni^{2+}	**nickelous**	**Ni^{3+}**	**nickelic**
Hg_2^{2+}	**mercurous****	**Hg^{2+}**	**mercuric**

{***notice that Hg_2^{2+} is a polyatomic ion***}**

c) Cations existing as polyatomic ions are just named.

NH_4^{1+} ➡ **ammonium**

2) <u>Name the Anion</u>:

a) Anions formed from <u>nonmetal atoms</u> are named using the "ide" suffix. ***The "ide" suffix always = negative charge.***

F^{1-}	**fluoride**	**Cl^{1-}**	**chloride**
O^{2-}	**oxide**	**S^{2-}**	**sulfide**
N^{3-}	**nitride**	**P^{3-}**	**phosphide**

b) Anions existing as polyatomic ions are just named

NO_3^{1-}	➡	**nitrate**	**["ate" = negative charge]**
$C_2H_3O_2^{1-}$	➡	**acetate**	**[CH_3COO^{1-}]**
ClO^{1-}	➡	**hypochlorite**	**[OCl^{1-}]**
ClO_2^{1-}	➡	**chlorite**	**["ite" = negative charge]**
ClO_3^{1-}	➡	**chlorate**	
ClO_4^{1-}	➡	**perchlorate**	
SO_4^{2-}	➡	**sulfate**	
CO_3^{2-}	➡	**carbonate**	
HCO_3^{1-}	➡	**hydrogen carbonate**	**[bicarbonate]**
MnO_4^{1-}	➡	**permanganate**	
CrO_4^{2-}	➡	**chromate**	
PO_4^{3-}	➡	**phosphate**	
OH^{1-}	➡	**hydroxide**	**["ide" = negative charge]**

<u>**Q1**</u>: Name the following ionic compounds:

a) MgO b) Cu_2S c) NH_4Cl d) $NaOCl$

<u>**A1**</u>: a) magnesium oxide

b) copper (I) sulfide or cuprous sulfide

c) ammonium chloride

d) sodium hypochlorite

B) Naming Inorganic Molecular Compounds:

1) Naming Binary Molecular Compounds:

a) Name the first element.

b) Name the second element with an "ide" ending.

c) Use prefixes with both names to denote number of atoms of each element in the formula. [mono, di, tri, tetra, penta, hexa, hepta, octa, nona, deca]

Exceptions:

(i) if there is only one atom of the ***first*** element no prefix is used. (ie omit the "mono") ex: *CO = carbon monoxide*

(ii) if the first element is hydrogen then prefixes are not used at all. ex: *HCl = hydrogen chloride*

Q2: Name the following binary molecular compounds:

a) $CS_{2(\ell)}$ b) $N_2O_{3(g)}$ c) $HI_{(g)}$ d) $PCl_{5(g)}$

A2: a) carbon disulfide

b) dinitrogen trioxide

c) hydrogen iodide

d) phosphorus pentachloride

FYI: There are a few important binary molecular compounds that are named with trivial names. These names need to be memorized.

NH_3 – ammonia, PH_3 – phosphine, N_2H_4 – hydrazine

2) Naming Acids (aqueous compounds beginning with H):

a) Acids whose anion name ends in "***ide***" are named by adding the prefix "***hydro***" to the anion name and then changing the anion's "***ide***" ending to "***ic***".

b) Acids whose anion name ends in "***ate***" are named by changing the anion's "***ate***" ending to "***ic***".

c) Acids whose anion name ends in "***ite***" are named by changing the anion's "***ite***" ending to "***ous***".

Q3: Name the following acids:

a) $H_2SO_{3(aq)}$ b) $HNO_{3(aq)}$ c) $HI_{(aq)}$ d) $H_2S_{(aq)}$

A3: a) $\boxed{\text{sulfurous acid}}$

b) $\boxed{\text{nitric acid}}$

c) $\boxed{\text{hydroiodic acid}}$

d) $\boxed{\text{hydrosulfuric acid}}$

VIII – Naming Organic Compounds

A) Hydrocarbons:

Although ***hydrocarbons*** are binary molecular compounds, composed of carbon and hydrogen, they are not named like the inorganic binary compounds. Instead, the names are often derived using prefixes that indicate the number of carbon atoms present and a root that indicates the homologous series to which the hydrocarbon belongs. The ***alkanes*** are hydrocarbon compounds that all have the same general formula $\mathbf{C_nH_{2n+2}}$.

The first ten members of the alkane homologous series are:

methane – CH_4; ethane – C_2H_6; propane – C_3H_8; butane – C_4H_{10};

pentane – C_5H_{12}; hexane – C_6H_{14}; heptane – C_7H_{16};

octane – C_8H_{18}; nonane – C_9H_{20}; decane – $C_{10}H_{22}$;

B) Alcohols:

Common alcohols are derivatives of alkanes in which a hydrogen atom has been replaced by the alcohol functional group "**OH**" called the hydroxyl group. To name such alcohols, replace the "e" of the alkane parent with an "ol". For alcohols with 3 or more carbon atoms, use a number to indicate on which carbon the "OH" is attached.

Some common monohydroxy alcohols are: methanol – CH_3OH;

ethanol – CH_3CH_2OH; 1-propanol – $CH_3CH_2CH_2OH$;

2-propanol – $CH_3CH(OH)CH_3$; 1-butanol – $CH_3CH_2CH_2CH_2OH$

C) Amines (Organic Bases):

Common amines (organic bases) are derivatives of alkanes in which a hydrogen atom has been replaced by the amine functional group "**NH_2**" called the amine group. To name such bases, replace the "ane" ending of the alkane parent with an "yl" and then add the word amine.

Some common amines (organic bases) are:
methyl amine – CH_3NH_2; ethyl amine – $C_2H_5NH_2$

Ammonia (NH_3) is not an organic compound but is a common base.

Identify each of the following compounds as an **acid**, a **base**, a **salt**, a **sugar**, an **alcohol**, a **hydrocarbon**, or a **binary oxide** … and then name the compound.

1) CO_2 - binary oxide & carbon dioxide

2) $C_6H_{12}O_6$ - sugar & glucose

3) CH_3OH - alcohol & methanol

4) $MgSO_4$ - salt & magnesium sulfate

5) $C_{12}H_{22}O_{11}$ - sugar & sucrose

6) CH_3COOH - acid & acetic acid

7) CH_3NH_2 - base & methyl amine

8) Na_2O - binary oxide & sodium oxide

9) KOH - base & potassium hydroxide

10) C_3H_7OH - alcohol & propanol

11) LiBr - salt & lithium bromide

12) $Ca(OH)_2$ - base & calcium hydroxide

13) HCl - acid & hydrochloric acid

14) C_6H_6 - hydrocarbon & benzene

15) MgO - binary oxide & magnesium oxide

16) C_2H_5OH - alcohol & ethanol

17) NH_3 - base & ammonia

18) NO_2 - binary oxide & nitrogen dioxide

19) $C_5H_{11}OH$ - alcohol & pentanol

20) NaOH - base & sodium hydroxide

21) $C_5H_{10}O_5$ - sugar & ribose

22) NH_4Cl - salt & ammonium chloride

23) NH_4NO_3 - salt & ammonium nitrate

24) C_5H_{12} - hydrocarbon & pentane

25) HIO_3 - acid & iodic acid

26) HNO_3 - acid & nitric acid

27) C_3H_8 - hydrocarbon & propane

28) $C_5H_{10}O_4$ - sugar & deoxyribose